SECTION B

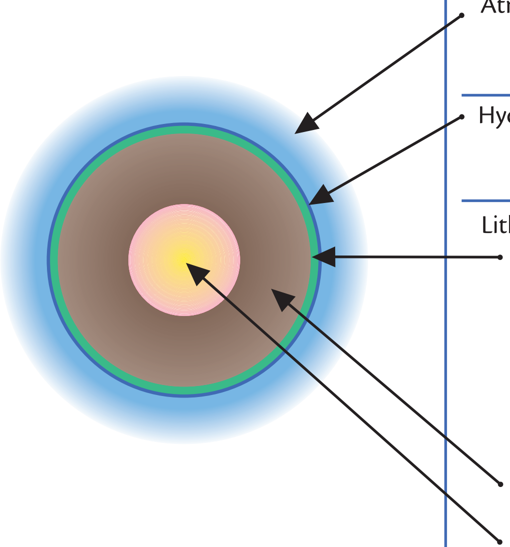
EARTH'S MINERAL RESOURCES

Among Earth's resources, metals—and the minerals from which they are extracted—have had long-standing importance for and use by humans. Those uses have ranged from toolmaking, energy transmission, and construction to works of art, decoration, and coin making. In this section, you will explore the properties and uses of minerals and metals. Using copper as a case study, you will learn about Earth's mineral resources and how they are converted to pure metals.

B.1 SOURCES AND USES OF METALS

Earth's Mineral Resources CD-ROM WWW.

Human needs for resources—whether to create a new coin, manufacture new clothing, construct a space-vehicle launch rocket, or supply fertilizer for food crops—must all be met by chemical supplies currently present on Earth. These supplies of resources are often cataloged by where they are found. The table in Figure 9 indicates the composition of Earth.



Earth's Composition

Layer of Planet	Thickness (Average)	Composition (Decreasing Order of Abundance)
Atmosphere	100 km	N_2 (78%), O_2 (21%), Ar (0.9%), He + Ne (<0.01%), variable amounts of H_2O, CO_2, etc
Hydrosphere	5 km	H_2O, and in ocean water approximately 3.5% NaCl and smaller amounts of Mg, S, Ca, and other elements as ions
Lithosphere: Crust	6400 km: Top 40 km	Silicates (compounds of metals, Si, and O atoms). Metals include Al, Na, Fe, Ca, Mg, K, and others Coal, oil, and natural gas Carbonates such as $CaCO_3$ Oxides such as Fe_2O_3 Sulfides such as PbS
Mantle	40–2900 km	Silicates of Mg and Fe
Core	2900 km to Earth's center	Fe and Ni

Figure 9 *Table of Earth's composition.*

Earth's atmosphere, hydrosphere, and outer layer of the lithosphere (the solid part of Earth) supply all resources for all human activities. The atmosphere provides nitrogen, oxygen, neon, argon, and a few other gases. From the hydrosphere come water and some dissolved minerals. The lithosphere provides the greatest variety of chemical resources. For example, petroleum and metal-bearing ores are found there. An **ore** is a naturally occurring rock or mineral that can be mined and from which it is profitable to extract a metal or other material. An ore contains a mixture of components. Of these, the most important are **minerals**—solid compounds containing the element or group of elements of interest.

The deepest mines on Earth barely scratch the surface of its crust. If Earth were the size of an apple, all accessible resources of the lithosphere would be located within the apple's skin. From this thin band of soil and rock, we obtain the major raw materials needed to build homes, automobiles, appliances, computers, videotapes, compact discs, and sports equipment—in fact, all manufactured objects.

As you can see from the table in Figure 10 on page 114, many of Earth's important resources are not uniformly distributed. There is no connection between a nation's supply of these resources and either its land area or its population. Quite often a particular region serves as the predominant supplier of certain metals vitally important to industry. For example, Africa holds much of the world's known reserves of chromium (80%), cobalt (54%), and manganese (61%).

The development of the United States as a major industrial nation has been facilitated, in part, because of the quantity and diversity of chemical resources found here. Yet in recent years, the United States has imported increasing amounts of certain vital chemical resources. For example, about 75% of the nation's tin (Sn) is imported.

The greatest challenge regarding mineral resources is deciding on the wisest uses of the available supplies. Immediate issues include some that have technical and economic implications. For example, is it worthwhile to mine a particular metallic ore at a certain site? The answer depends on several factors:

- amount of useful ore at the site
- percent of metal in the ore
- type of mining and processing needed to extract the metal from its ore
- distance of the mine from metal refining facility and markets
- metal's supply-versus-demand status

Copper, one of the materials you might be thinking of using in your coin design, provides a case study of a vital chemical resource. You will first consider worldwide sources of copper and how these copper-bearing materials are converted to pure copper. Later, you will explore some possible replacements for this resource.

Copper is one of the most familiar and widely used metals in modern society. Among all the elements, it is second only to silver in electrical conductivity. This property and its relatively low cost, corrosion resistance, and ductility (ease of being drawn into thin wires), make copper the world's most common metal for electrical wiring. Copper is also used to produce

Worldwide Annual Production of Selected Metals					
Metal	Country	Percent production	Actual production (1000 metric tons)	World total production (1000 metric tons)	Year
Aluminum	United States	17%	3713	22 100	1998
	Russia	14%	3005		
	Canada	11%	2374		
	China	10%	2100		
	Australia	7%	1627		
Copper†	United States	15%	1720	11 100	1997
	Chile	13%	1389		
	Japan	12%	1350		
	China	9%	963		
	Russia	5%	600		
Iron ore○	United States	22%	65 900	305 300	1997
	Brazil	12%	37 300		
	Russia	11%	34 000		
	Ukraine	10%	32 000		
	Canada	9%	27 300		
Lead□	United States	22%	1450	5880	1998
	United Kingdom	6%	350		
	Germany	6%	335		
	France	5%	306		
	Japan	5%	302		
Nickel△	Russia	23%	260	1120	1997
	Canada	17%	190		
	New Caledonia	12%	137		
	Australia	11%	124		
	Indonesia	6%	72		
Silver△	Mexico	16%	2.7	16.4	1998
	United States	13%	2.1		
	Peru	12%	1.9		
	Australia	9%	1.5		
	China	9%	1.4		
Tin†	China	29%	61	213	1997
	Indonesia	19%	40		
	Malaysia	17%	36		
	Brazil	9%	19		
	Bolivia	8%	16		
Zinc†	China	18%	1500	3890	1997
	Canada	9%	743		
	Japan	8%	653		
	Republic of Korea	5%	390		
	Spain	5%	370		

† = world smelter production ○ = world pelletizing capacity □ = world refinery production △ = world mine production

Figure 10 *Production of selected metals worldwide.*

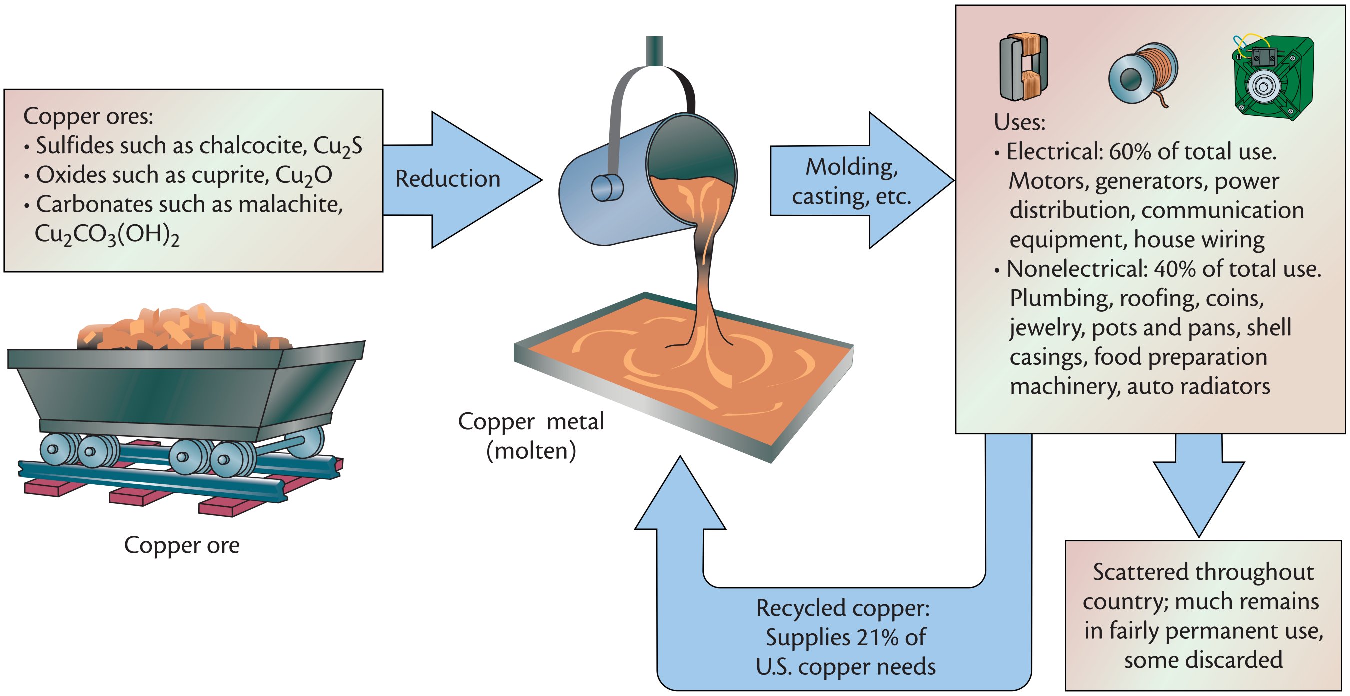


Figure 11 *The copper cycle and copper uses.*

brass, bronze, and other alloys, a variety of important copper-based compounds, jewelry, and works of art.

Alloys are discussed in more detail in Section D.

The first copper ores mined contained from 35% to 88% copper. Although such ores are no longer available, ores less rich in copper can be used. In fact, it is now economically possible to mine ores containing less than 1% copper. Copper ore is chemically processed to produce metallic copper, which is then transformed into a variety of useful materials. Figure 11 summarizes the copper cycle from sources to common uses to waste products.

Earth's accessible deposits of this valuable resource are destined to be depleted. Will future developments increase or decrease the need for copper? What copper substitutes are available? The following activity will help you address these questions.

PROPERTIES AND USES OF COPPER

Building Skills 4

Some of copper's properties are listed in the table in Figure 12 (page 116). Consider how these properties make copper suitable for uses depicted in Figure 11. For example, what properties make copper useful in electrical power generators? Copper's high electrical conductivity is certainly essential to this application. Copper's malleability and ductility are also important, making it possible to form copper wires and to wrap them in a generator. Corrosion resistance is also a benefit in such large, expensive equipment or in other applications.

Properties of Copper	
Malleability and ductility	High
Electrical conductivity	High
Thermal conductivity	High
Chemical reactivity	Relatively low
Resistance to corrosion	High
Useful alloys formed	Bronze, brass, etc.
Color and luster	Reddish, shiny

Figure 12 *Properties of copper.*

1. Consider the remaining uses of copper listed in Figure 11. For each use, identify those particular properties that make copper an appropriate choice.
2. a. How would increased recycling of scrap copper affect future availability of this metal?
   b. Is there a limit to the role copper recycling can play? Why?
3. For each use listed below, describe a technological change that could decrease the demand for copper:
   a. coins
   b. communications
   c. power generation
   d. indoor electrical wiring

B.2 CONVERTING COPPER

Laboratory Activity

Introduction

You have seen many chemical reactions in your lifetime. Some, such as a fireworks display, are memorable. Others, such as the slow process of rusting, are far less dramatic. Have you ever stopped to think about what happens to the atoms involved in those reactions? Are the materials that made up the fireworks still there after they are launched into the sky and ignited? What about the iron that turns into rust?

In this laboratory activity, you will work with a powdered sample of elemental copper, a metal you may be considering for your coin design. As you observe its chemical behavior, think about whether its properties make it a good candidate for this use.

Procedure

1. Prepare a data table to record the masses you will determine in Steps 2 and 9.
2. Measure and record the mass of a clean, empty crucible. Add approximately 1 g copper powder to the crucible. Record the mass of the crucible with copper powder in it within the nearest 0.1 g. Find

the actual mass of copper powder by subtracting the mass of the empty crucible from this value. Record the mass of copper powder.

3. Which properties of copper can be directly observed? Record your observations of the copper powder.
4. Set up the crucible, clay triangle, and burner as shown in Figure 13. The crucible lid should be slightly ajar.
5. Light the burner, and adjust it so that the flame tip just touches the bottom of the crucible.
6. Heat the crucible and its contents for two minutes. Remove the flame, and use a spatula to break up the solid in the crucible so that as much remaining copper metal is exposed as possible.
   CAUTION: *Avoid touching the hot crucible.*
7. Continue heating for about 10 minutes more, removing the flame and breaking up the solid every two to three minutes.
8. When you have finished the heating, extinguish the burner flame and allow the crucible and its contents to cool to room temperature. Answer Questions 1 and 2 while you are waiting.
9. After the crucible and its contents have cooled, find their mass. Use this value and the mass of the empty crucible to calculate the mass of the contents. Record these values in your data table.
10. Transfer your product to a clean 100-mL beaker. Label and store the beaker and product as indicated by your teacher.
11. Put away the other materials. Wash your hands thoroughly before leaving the laboratory.

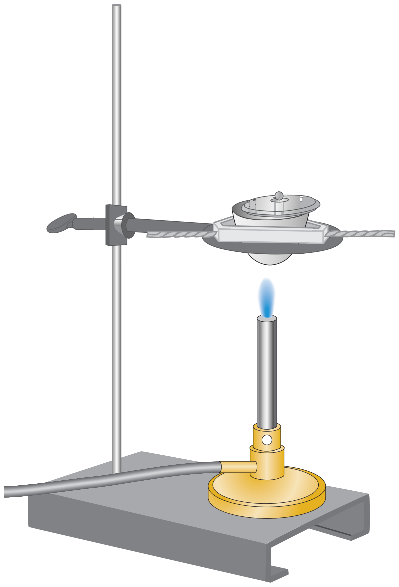


Figure 13 *Clay triangle holding a crucible over a burner.*

Questions

1. a. Were the changes you observed physical or chemical?
   b. How do you know?
2. a. Describe the changes you observed as you heated the copper.
   b. Did the copper atoms remain in the crucible? Explain.
3. a. What happened to the mass of the crucible contents after you heated the copper?
   b. Why do you think the mass changed?

B.3 METAL REACTIVITY

As you just observed, when copper metal is heated, it gradually reacts with oxygen in the air to produce a black substance. This is the equation for the reaction:

$$\underset{\text{Copper}}{2\,Cu(s)} + \underset{\text{Oxygen}}{O_2(g)} \longrightarrow \underset{\text{Copper(II) oxide}}{2\,CuO(s)}$$

Although copper reacts to form copper(II) oxide when heated, at room temperature the metal remains relatively unreactive in air. You are probably familiar with this fact from observing that copper wire and the copper surface on pennies do not turn black under normal conditions.

There are two common compounds of copper and oxygen—CuO and Cu_2O. Because the name "copper oxide" could be applied to both, a Roman numeral is added to indicate copper's ionic charge. Copper(I) oxide is Cu_2O because it contains Cu^+ ions; copper(II) oxide is CuO because it contains Cu^{2+} ions.

Figure 14 *Magnesium metal and oxygen gas react so spectacularly that small samples of magnesium are used in some fireworks.*

Magnesium metal also reacts with oxygen gas. But unlike copper metal, magnesium heated in air ignites and produces a brief, blinding flash. See Figure 14. This is the equation for the reaction:

$$\underset{\text{Magnesium}}{2\ \mathrm{Mg(s)}} + \underset{\text{Oxygen}}{\mathrm{O_2(g)}} \longrightarrow \underset{\text{Magnesium oxide}}{2\ \mathrm{MgO(s)}}$$

By contrast, gold (Au) does not react with any components in air; in particular, it does not react with oxygen gas, even at elevated temperatures. This is one reason why gold metal is highly prized in long-lasting, decorative objects, such as jewelry. Gold-plated electrical contacts, such as those used for automobile air bags and audio cable connectors, are very dependable because nonconducting oxides do not form on the contact surfaces.

Observing how readily a certain metal reacts with oxygen provides information about the metal's reactivity. If elements are ranked in relative order of their chemical reactivity, the ranking is called an **activity series**. Based on what you have just learned about gold and magnesium and what you already know about copper, how would you rank the three metals in terms of their relative chemical reactivity?

B.4 RELATIVE REACTIVITIES OF METALS

Laboratory Activity

Introduction

In this activity, you will investigate the reactions of the metals copper, magnesium, and zinc with solutions that each contain a metal cation. The four solutions you will use are copper(II) nitrate, $Cu(NO_3)_2$ (containing Cu^{2+});

magnesium nitrate, $Mg(NO_3)_2$ (containing Mg^{2+}); zinc nitrate, $Zn(NO_3)_2$ (containing Zn^{2+}); and silver nitrate, $AgNO_3$ (containing Ag^+).

Procedure

1. Devise an orderly procedure that will allow you to observe the reaction (if any) between each metal and each of the four ionic solutions. You will conduct each reaction in a separate well of your wellplate, using five drops of 0.2 M solution and a small strip of metal. How many different combinations of metals and solutions will you need to observe? How will you arrange things so you can complete your observations efficiently yet remain certain which metal and which solution are in each well?
2. Prepare a data table to help you organize the observations and results of the procedure you devise.
3. Obtain 5-mm strips of each of the three metals to be tested. Clean the surface of each metal strip by rubbing it with sandpaper or emery paper. Record observations of each metal's appearance.
4. Complete your planned procedure, writing your observations in your data table. **CAUTION:** *Avoid letting the $AgNO_3$ solution come in contact with skin or clothing as it causes dark, non-washable stains.* If no reaction is observed, write NR in the table. Record the observed changes if a reaction occurs.
5. Dispose of your solid samples and wellplate solutions as directed by your teacher.
6. Wash your hands thoroughly before leaving the laboratory.

Questions

1. Which metal reacted with the most solutions?
2. Which metal reacted with the fewest solutions?
3. With which of the solutions (if any) would you expect silver metal to react, if it were available to be tested?
4. List the metals (including silver) in order, placing the most reactive metal first (the one reacting with the most solutions) and the least reactive metal last (the one reacting with the fewest solutions).
5. Refer to your "metal activity series" list in Question 4. Write a brief explanation of why the outside surface of a penny is made of copper instead of zinc.
6. a. Which of the four metals mentioned in this laboratory activity might be an even better choice than copper for the outside surface of a penny? Why?
   b. Why do you think that metal is not used for that purpose?
7. Given your new knowledge about the relative chemical activities of these four metals,
   a. which metal is most likely to be found in an uncombined, or "free," (metallic) state in nature?
   b. which metal is least likely to be found chemically uncombined with other elements?

Relative Reactivities of Metals CD-ROM WWW.

8. Reconsider your experimental design for this activity.
   a. Would it have been possible to eliminate one or more of the metal-solution combinations and still obtain all information needed to create chemical activity ratings for the metals?
   b. If so, which combination(s) could have been eliminated? Why?

Figure 15 *These stone, bronze, and iron tools represent three major ages of civilization.*

B.5 METALS: PROPERTIES AND USES

Humans have been described as toolmakers. Readily available stone, wood, and natural fibers were the earliest materials used to make useful tools—hammers, chisels, knives, spears, and grinding devices. It was the discovery that fire could transform materials in certain rocks into strong, malleable metals, however, that triggered a dramatic leap in the growth of civilization.

Gold and silver, found as free elements rather than in chemical combination with other elements, were probably the first metals used by humans. These metals were formed into decorative objects and, later, into coins. Their relative unreactivity made them excellent materials for those uses.

It is estimated that copper has been used to make tools, weapons, utensils, and decorations for about 10 000 years. Bronze, an alloy of copper and tin, was developed about 3800 B.C. Thus humans moved from the Stone Age into the Bronze Age. See Figure 15.

Eventually early people developed iron metallurgy, the extraction of iron from its ores. This led to the start of the Iron Age, more than 3000 years ago. In time, as humans learned more about chemistry and fire, a variety of metallic ores were transformed into increasingly useful metals.

DISCOVERY OF METALS — ChemQuandary 1

Copper, gold, and silver are far from being the most abundant metals on Earth. Aluminum, iron, and calcium, for example, are all much more plentiful. Why, then, were copper, gold, and silver among the first metallic elements discovered?

You have explored some of the chemistry of metals and know, for example, that copper metal is more reactive than silver but less reactive than magnesium. A more complete activity series is given in the table in Figure 16. The table also includes brief descriptions of common methods for retrieving each metal from its ore.

In this list the most reactive metallic elements are at the top; less reactive elements are closer to the bottom. An activity list can be used to predict whether certain reactions can be expected. For example, you observed in the laboratory that zinc metal, which is more reactive than copper, reacted with copper ions in solution. However, zinc metal did not react with magnesium ions in solution. Why? Because zinc is less reactive than magnesium.

Metal Activity Series			
Element	Metal Ion(s) Found in Minerals	Process Used to Obtain the Metal	State of Metal Obtained
Lithium	Li^{+}	Pass direct electric current through the molten mineral salt (electrometallurgy)	Li(s)
Potassium	K^{+}		K(s)
Calcium	Ca^{2+}		Ca(s)
Sodium	Na^{+}		Na(s)
Magnesium	Mg^{2+}		Mg(s)
Aluminum	Al^{3+}		Al(s)
Manganese	Mn^{2+}	Heat mineral with coke (C) or carbon monoxide (CO) (pyrometallurgy)	Mn(s)
Zinc	Zn^{2+}		Zn(s)
Chromium	Cr^{3+}, Cr^{2+}		Cr(s)
Iron	Fe^{3+}, Fe^{2+}		Fe(s)
Lead	Pb^{2+}	Heat (roast) mineral in air (pyrometallurgy) or find the element free (uncombined)	Pb(s)
Copper	Cu^{2+}, Cu^{+}		Cu(s)
Mercury	Hg^{2+}		Hg(l)
Silver	Ag^{+}		Ag(s)
Platinum	Pt^{2+}		Pt(s)
Gold	Au^{3+}, Au^{+}		Au(s)

Figure 16 *Metal activity series (in decreasing order of reactivity).*

In general, a more reactive metallic element (higher in the activity series) will cause ions of a less reactive metallic element (lower in the activity series) to change to their corresponding metal.

TRENDS IN METAL ACTIVITY

Building Skills 5

Use the table in Figure 16 and the Periodic Table (page 104) to answer the following questions.

1. a. What trend in metallic reactivity is found from left to right across a horizontal row (period) of the Periodic Table? (*Hint:* Compare the reactivities of sodium, magnesium, and aluminum.)
   b. In which part of the Periodic Table are the most-reactive metals found?
   c. Which part of the Periodic Table contains the least-reactive metals?

2. a. Will iron (Fe) metal react with a solution of lead(II) nitrate, $Pb(NO_3)_2$?
   b. Will platinum (Pt) metal react with a lead(II) nitrate solution?
   c. Explain your answers to Questions 2a and 2b.

3. Use specific examples from the activity series in your answers to these two questions:
   a. Are least-reactive metals also the cheapest metals?
   b. If not, what other factor(s) might influence the market value of a metal?

B.6 MINING AND REFINING

The process of converting a combined metal (usually a metal ion) in a mineral to a free metal involves a particular kind of chemical change. For example, the conversion of a copper(II) cation to an atom of copper metal requires two electrons.

Formation of Copper Metal

$$\underset{\text{Copper(II) ion}}{Cu^{2+}} + 2\,e^{-} \longrightarrow \underset{\text{Copper metal}}{Cu}$$

In general, to convert metal cations to neutral metal atoms, each cation must gain one or more electrons.

Chemists classify any chemical change in which a species gains one or more electrons as a **reduction.** Thus the conversion of copper(II) cations to copper metal is a reduction reaction. You can convince yourself of this fact by examining the equation above for that change. Chemists classify the reverse reaction, in which an ion or other species loses one or more electrons, as an **oxidation.** For example, under the right conditions copper atoms can be oxidized.

Formation of Copper(II) Ions

$$\underset{\text{Copper metal}}{Cu} \longrightarrow \underset{\text{Copper(II) ion}}{Cu^{2+}} + 2\,e^{-}$$

Historically, "oxidation" referred to the chemical combination of a substance with oxygen, as the term itself suggests. Chemists now know that in nearly all cases in which oxygen combines with another element or compound, oxygen partially or fully removes one or more electrons from the other species. By today's definition, any reactant—be it oxygen or not—that causes a species to lose one or more electrons is said to cause that species to be oxidized.

One way to remember this is **OIL RIG: O**xidation **I**s **L**oss of electrons, **R**eduction **I**s **G**ain of electrons.

Whenever one species loses electrons, another species must simultaneously gain them. In other words, oxidation and reduction reactions never occur separately. Oxidation and reduction occur together in what chemists call **oxidation-reduction reactions** or, to use a common chemical nickname, **redox reactions.**

You have already observed redox reactions in the laboratory. In Laboratory Activity B.4 (pages 118–120), copper metal reacted with silver ions. Here is the oxidation-reduction reaction you observed:

$$\underset{\text{Copper metal}}{Cu(s)} + \underset{\text{Silver ion}}{2\,Ag^{+}(aq)} \longrightarrow \underset{\text{Copper(II) ion}}{Cu^{2+}(aq)} + \underset{\text{Silver metal}}{2\,Ag(s)}$$

A Copper Redox Reaction CD-ROM WWW.

Each metallic copper atom (Cu) was oxidized (converted to a Cu^{2+} ion by losing two electrons) and each silver ion (Ag^{+} from $AgNO_3$ solution) was reduced (converted to an Ag atom by gaining one electron).

In the same activity you found that copper ions could be recovered from solution as copper metal by allowing the copper ions to react with magnesium metal, an element more active than copper. Magnesium atoms were oxidized; copper ions were reduced. Do you see why?

Note that the total electrical charge on both sides of this equation is the same. Electrical charges—as well as atoms—must balance in a correctly written chemical equation.

$$\underset{\text{Copper(II) ion}}{Cu^{2+}(aq)} + \underset{\text{Magnesium metal}}{Mg(s)} \longrightarrow \underset{\text{Copper metal}}{Cu(s)} + \underset{\text{Magnesium ion}}{Mg^{2+}(aq)}$$

In some circumstances this reaction might be a useful way to obtain copper metal. However, as is often the case, the desired copper metal is gained at the expense of "using up" another highly desirable material—in this case, magnesium metal.

How do redox reactions occur? Many metallic elements are found in minerals in the form of cations because they combine readily with other elements to form ionic compounds. Obtaining a metal from its mineral requires energy and a source of electrons. A reacting chemical species that serves as the source of electrons is known as a **reducing agent**.

Look again at Figure 16 (page 121). The table highlights several techniques that are used to reduce metal cations—or, in other words, to supply one or more electrons to each cation. The specific technique chosen depends on the metal's reactivity and the availability of inexpensive reducing agents and energy sources.

Two major approaches summarized in the table are **electrometallurgy** and **pyrometallurgy.** As the table suggests, electrometallurgy involves using an electric current to supply electrons to metal ions, thus reducing them. This process is used when no adequate chemical reducing agents are available or when very high-purity metal is sought. Pyrometallurgy—the most important and oldest ore-processing method—involves the treatment of metals and their ores by heat, as in a blast furnace. Carbon (coke) and carbon monoxide are common reducing agents in pyrometallurgy. A more active metal can be used if neither of these will do the job.

A third approach to obtaining metals from their ions is the process called **hydrometallurgy**—the treatment of ores and other metal-containing materials by reactants in water solution. You used such a procedure when you investigated the reactivity of different metals in Laboratory Activity B.4. Hydrometallurgy is used to recover silver and gold from old mine tailings (the mined rock left after most of the sought mineral is removed) by a process known as leaching. As supplies of higher-grade ores become scarcer, it will become economically feasible to use hydrometallurgy and other "wet processes" on metal-bearing minerals that dissolve in water.

MODELING MATTER

ELECTRONS AND REDOX PROCESSES

The processes of oxidation (loss of one or more electrons) and reduction (gain of one or more electrons) can be clarified by visual representations of the events. To develop such representations, you will consider atoms of each of the metals you investigated in Laboratory Activity B.4.

First, however, a review of some key details about the composition of an atom is in order. Magnesium (Mg), an active metal, formed Mg^{2+} ions in several of the reactions. The atomic number of Mg is 12, indicating that an electrically neutral atom of magnesium contains 12 protons and 12 electrons. (Do you recall why those numbers must be equal for a neutral atom?)

If magnesium forms a Mg^{2+} ion, two negatively charged electrons must be removed from each magnesium atom. The bookkeeping involved in this change can be summarized this way:

Mg	$\longrightarrow$	Mg^{2+}	+	$2\,e^-$
12 protons (+)		12 protons (+)		
12 electrons (−)		10 electrons (−)		2 electrons (−)
Net charge: 0		Net charge: 2+		Net charge: 2−

To build a useful picture of this process in your mind, it is necessary to keep track of only the two electrons that each magnesium atom releases, rather than monitoring all 12 of the atom's available electrons. (In fact, in normal chemical reactions, a magnesium atom is not observed to release any of its other 10 electrons.)

Thus, for bookkeeping purposes, an atom of Mg will be depicted this way:

Mg: the symbol for the element with two dots attached.

Each dot represents one readily removable electron. The symbol Mg represents the remaining parts of a magnesium atom, including its other ten electrons. The resulting expression for Mg is called an **electron-dot structure,** or just a **dot structure.** The equation for the oxidation of Mg can be represented in electron-dot terms this way:

$$\text{Mg:} \longrightarrow Mg^{2+} + 2\,e^-$$

1. Construct a similar electron-dot expression for the change that occurred in Laboratory Activity B.4 when each of these events took place:
   a. An atom of zinc, Zn, was converted to a Zn^{2+} ion. (*Hint:* Zn has two readily removable electrons.)
   b. A silver ion, Ag^+, was converted to a metallic silver atom, Ag(s).
2. Apply the definitions of oxidation and reduction to your two equations in Question 1, and label each reaction appropriately.

Now consider one of the complete reactions that you observed in Laboratory Activity B.4. When you

immersed a sample of copper metal, Cu, in silver nitrate solution, $AgNO_3$, a blue solution containing Cu^{2+} formed, as well as crystals of solid Ag. This is the reaction that occurred:

$$Cu(s) \quad + \quad 2\,Ag^{+}(aq) \quad \longrightarrow \quad Cu^{2+}(aq) \quad + \quad 2\,Ag(s)$$

Using dot structures, the reaction can be represented this way:

$$Cu\!: \; + \; Ag^{+} \; + \; Ag^{+} \; \longrightarrow \; Cu^{2+} \; + \; Ag\cdot \; + \; Ag\cdot$$

3. a. Which reactant (Cu or Ag^{+}) is oxidized?
   b. Which is reduced?
4. Why are two Ag^{+} ions needed for each Cu(s) atom that reacts?

Each copper atom involved in this reaction loses two electrons. Thus copper atoms must be oxidized in the change. It is clear from the dot structures that those two electrons lost by copper are gained by two Ag^{+} ions. So Ag^{+} is the agent that caused the removal of electrons from Cu (resulting in the oxidation of Cu). The species involved in removing electrons from the reactant that is oxidized is called the **oxidizing agent**—in this case, Ag^{+} ions.

5. a. Given that definition and explanation, what must be the reducing agent in the reaction between Cu(s) and Ag^{+} ions?
   b. How would you define a reducing agent?

Now consider another reaction you observed in Laboratory Activity B.4:

$$Zn(s) \quad + \quad Cu^{2+}(aq) \quad \longrightarrow \quad Zn^{2+}(aq) \quad + \quad Cu(s)$$

6. Draw an electron-dot representation of this reaction.
7. a. Which reactant is oxidized?
   b. Which is reduced?
8. Identify the oxidizing agent and the reducing agent in this reaction.
9. Consider both of the oxidation-reduction reactions you analyzed in this exercise. What general features of an oxidation-reduction reaction would allow you to answer Questions 7 and 8 *without* drawing electron-dot representations?
10. Test your answer to Question 9 by considering a new oxidation-reduction reaction. Answer Questions 7 and 8 for this system:

$$Zn^{2+}(aq) \quad + \quad Mg(s) \quad \longrightarrow \quad Zn(s) \quad + \quad Mg^{2+}(aq)$$

SECTION SUMMARY

Reviewing the Concepts

◆ **The resources for all human activities must be obtained from Earth's atmosphere, hydrosphere, and outer layer of its lithosphere. These resources are not uniformly distributed.**

1. a. List and briefly describe the three major "parts" of Earth.
   b. Which part serves as the main storehouse of chemical resources used in manufacturing consumer products?
2. List two resources typically found in each of the three major parts of Earth.
3. According to information in Figure 10 on page 114, which of these four countries—the United States, Australia, China, or Brazil—produces the largest mass of these eight resources?
4. Is there a connection between the distribution of mineral resources and the wealth of a nation? Explain.

◆ **The feasible mining and extraction of a mineral resource depends, in part, on the amount of the resource available and the total cost of processing.**

5. What factors determine the feasibility of mining a particular metallic ore at a certain site?
6. A nineteenth-century gold mine, inactive for a hundred years, has recently reopened for further mining. What factors may have influenced the decision to reopen the mine?
7. What is meant by referring to the amount of "useful ore" at a site?

◆ **The ease with which a particular metal may be processed and preserved depends on its chemical reactivity. Active metals are more difficult to process than less-active metals and tend to corrode more quickly.**

8. Why are active metals more difficult to process and refine?
9. Based on your results from Laboratory Activity B.4, which metals involved in this activity would be the easiest to process? Why?
10. Why do most metals exist in nature as minerals rather than as pure metallic elements?
11. Consider these two equations. Which represents a reaction that is more likely to occur? Why?
    a. $Zn^{2+}(aq) + 2\ Ag(s) \longrightarrow Zn(s) + 2\ Ag^{+}(aq)$
    b. $2\ Ag^{+}(aq) + Zn(s) \longrightarrow 2\ Ag(s) + Zn^{2+}(aq)$
12. a. Why would it be a poor idea to stir a solution of lead(II) nitrate with an iron spoon?
    b. Write a chemical equation to support your answer.

◆ **The processes of oxidation (the loss of electrons) and reduction (the gain of electrons) occur together, resulting in oxidation-reduction (redox) reactions.**

13. Write an equation for each of these processes.
    a. the reduction of gold(III) ions
    b. the oxidation of elemental vanadium to vanadium(II) ions
    c. the oxidation of magnesium metal
14. Identify each equation as representing either an oxidation or a reduction reaction.
    a. $Fe^{2+} + 2\ e^{-} \rightarrow Fe$
    b. $Cr \rightarrow Cr^{3+} + 3\ e^{-}$
    c. $Al^{3+} + 3\ e^{-} \rightarrow Al$
15. Consider the following equation:
    $$Zn(s) + Ni^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Ni(s)$$
    a. Which reactant has been oxidized? Explain your choice.
    b. Which reactant has been reduced? Explain your choice.
    c. What is the reducing agent in this reaction?

◆ **Metal cations can be converted to metal atoms by electrometallurgy, pyrometallurgy, or hydrometallurgy.**

16. Explain how each process converts metal cations to metal atoms.
    a. electrometallurgy
    b. pyrometallurgy
    c. hydrometallurgy

17. What processes could you use to obtain these elements from their ores?
    a. magnesium
    b. lead

Connecting the Concepts

18. How can a less active metal be used to prevent the corrosion of a more active metal?
19. Large gold nuggets with masses of 45 kg (100 pounds) or more have been discovered. What conditions might have allowed such large pieces of elemental gold to exist?
20. There are thousands of tons of gold in sea water. Explain why it is unlikely that ocean water will ever be "mined" for gold.
21. In 1982, when the penny was converted from pure copper to copper and zinc, the outside surface was still coated with copper. List three reasons why this coating was used.
22. At one time, food cans were made with tin and soldered with lead. What kinds of health hazards were posed by this arrangement?
23. Is there any connection between the process used to reduce a metal cation and the position of that element on the Periodic Table?

Extending the Concepts

24. Although aluminum is a more reactive metal than iron, it is often used for outdoor products. Investigate how this is possible.
25. The uneven distribution of mineral resources sometimes affects relations between nations. Identify and describe one historical or current example of this fact.
26. What conclusions about materials can be drawn from a study of the substances used for currency in ancient civilizations? Explain your ideas by giving examples.
27. History documents that copper has been used by humans for 10 000 years, whereas aluminum has been used for only about 100 years. Suggest and explain some reasons for this difference.
28. The reactive metal aluminum is often used in containers for acidic beverages. Investigate and describe the technology that makes this possible.
29. What is a patina? Explain its value both aesthetically and chemically.